Chapter 14 – Chemical Periodicity and Trends

Watch the **Professor Dave Explains** video, *The Periodic Table: Atomic Radius, Ionization Energy, and Electronegativity* at https://www.youtube.com/watch?v=hePb00CqvP0

Arrangement of the Periodic Table

1. Why was Dimitri Mendeleev's model for the period table more powerful than other models proposed during the mid-to-late 1800's?

Mendeleev's table correlated data well, by organizing the periodic table in both vertical columns and horizontal rows; it also enabled much predictive power, since all elements in a group had similar properties, as did those in rows.

2. Elements arranged in columns are called _Groups/ Families___, and those arranged in rows are called _Periods_.

3. The arrangement of the table in this manner, enabled chemists to predict the properties of elements never before discovered, which represented gaps in the table. **True** or false? Circle one.

4. Why do elements in the same group (such as Li, Na, K) all behave similarly?

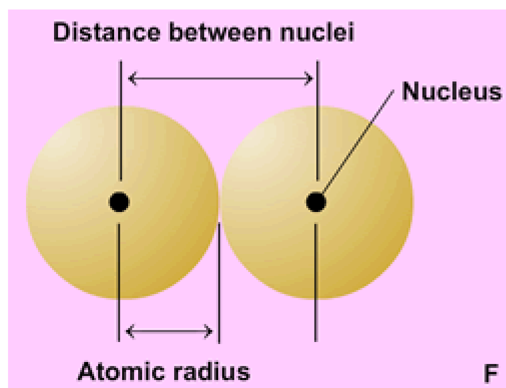
Elements within the same group all have the same electronic configuration with respect to the outermost energy level. As an example, Li, Na, and K, all have an ns^1 valence electron, where n = 1, n = 2, and n =3, respectively.

5. Every element in Group II, has how many electrons? ___2__ Name three of those elements:

(1)__beryllium___, (2) __magnesium__, (3)__calcium__

Periodic Trends in Atomic Radius

The atomic radius is one-half the distance between the nuclei of two like atoms in a real or theoretical diatomic molecule (Wilbraham, Staley, Matta, & Waterman, 2002, p. 398).



6. Describe the trend in atomic radius as you move down a group; as you move from hydrogen, to lithium, to sodium, etc.

As you move down a group, the atomic radius gets larger, because principal energy shells are added, to accommodate increasing numbers of electrons.

7. Why does this trend exist?

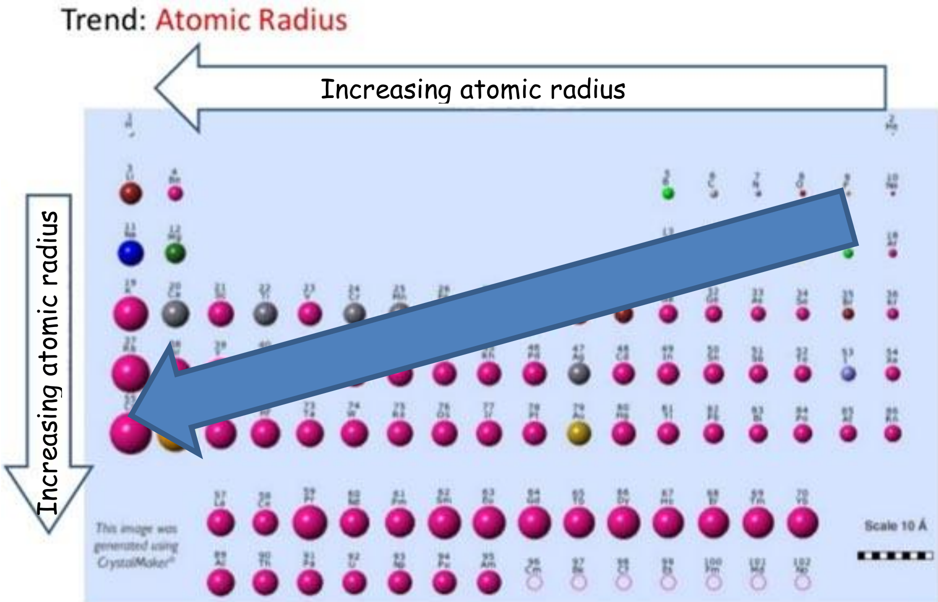
Energy shells, or principal quantum numbers are added as you move down a group. As an example, for H: $n = 1$; for Li: $n = 2$; for Na: $n = 3$.

8. Describe the trend in atomic radius as you move across a period; from sodium to magnesium, to aluminum, silicon and phosphorus.

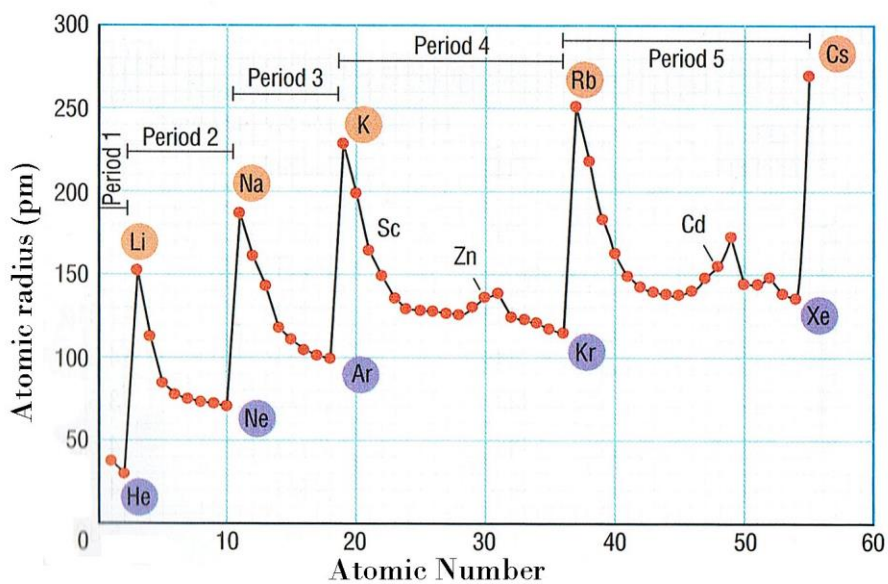
As you move across a period, from sodium to chlorine, atomic radius gets smaller.

9. What explains this trend? What role does the increasing number of protons play?

As the number of protons is increased in the nucleus, the pull on electrons is tighter, making the radius smaller.



10. Write in the arrows above, the trend of atomic size along a period and group, using the words *increasing* or *decreasing*. Then, write a single arrow on the table summarizing this trend.



11. Summarize the spikes in the graph plotting atomic number (x) verses atomic radius (y).

The spikes occur as the result of the addition of another electron shell; as the principal quantum number (*n*) increases, so does the atomic radius.

Periodic Trends in Ionic Radius

12. When an electron is removed from an atom, the radius gets __smaller__; when an electron is added to an atom, the radius gets __larger__.

13. What is the electron configuration of these ions (and neon atom)? O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+}

Hint: the electron configuration is the same for all these. __$1s^2 2s^2 2p^6$____

14. What accounts for the trend in ionic radius?

When the two valence electrons are removed from Mg, making it Mg^{2+}, the outermost energy level is no longer n = 3, but then becomes n = 2. When two electrons are added to O, making it O^{2-}, the ionic radius increases because the added repulsion forces makes the valence electrons move farther apart from one another, increasing the atomic radius.

Periodic Trends in Ionization Energy

The energy required to overcome the attraction of the nuclear charge and remove an electron from a gaseous atom is called the ionization energy. Removing one electron results in the formation of a positive ion with a 1+ charge (Wilbraham et al., 2002, p. 401).

$$Na(g) \rightarrow Na^{+}(g) + e^{-}$$

The energy required to remove the first outermost electron is called the *first ionization energy*.

15. What happens to the electromagnetic force attracting electrons to protons as the distance between outer electrons and the nucleus increases?

The attraction between the outermost electrons to that of the nucleus is reduced as the distance from the nucleus to the electrons is increased.

16. As you move down a group, from lithium to cesium, first ionization energy _decreases_, and as you move across a row, from sodium to argon, the first ionization energy _increases_.

17. Why is a francium outermost electron so much easier to pull off than an outermost electron from helium?

The outermost electron on Fr is much farther from the nucleus (n = 7) than that of He (n = 2), making it easier to pull off. In addition, when Fr loses the electron, the next energy level (n = 6) has a full energy shell, enabling some stability.

18. Why is the second ionization energy typically higher than the first?

The more electrons that are removed from an atom, the less stable it is.

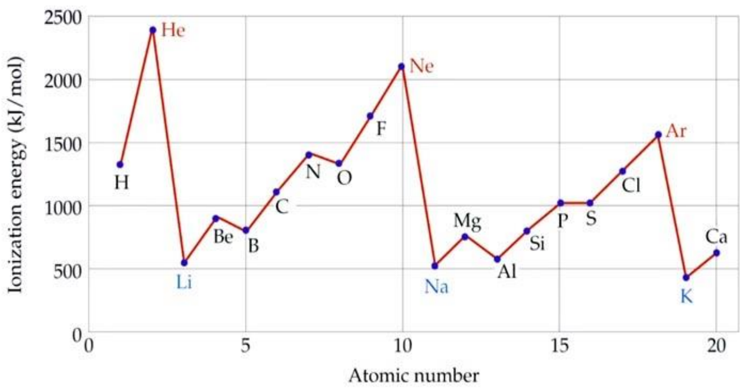
19. Once all the electrons are removed from an outer shell, there is a huge jump in ionization energy.

As an example, once the two outer electrons from Mg are removed, ionization energy goes from 1,451 kJ/ mol to 7,733 kJ/ mol. Why?

When electrons are removed from the outermost shell, the shell underneath is more resistant to give up any additional electrons.

successive ionization energies (kJ/mol)

	Na	Mg	Al	Si	P	S	Cl	Ar
1st IE	496	738	578	787	1,012	1,000	1,251	1,520
2nd IE	4,562	1,451	1,817	1,577	1,903	2,251	2,297	2,665
3rd IE	6,912	7,733	2,745	3,231	2,912	3,361	3,822	3,931
4th IE	9,543	10,540	11,575	4,356	4,956	4,564	5,158	5,770
5th IE	13,353	13,630	14,830	16,091	6,273	7,013	6,540	7,238
6th IE	16,610	17,995	18,376	19,784	22,233	8,495	9,458	8,781
7th IE	20,114	21,703	23,293	23,783	25,397	27,106	11,020	11,995



20. Note the graph at left. There is a trend in ionization energy that is obvious, with a few deviations within periods. Explain why it requires less ionization energy to pull an electron off of oxygen, than it does off of nitrogen?

Orbital symmetry: N's 3 p orbitals are ½ full making it more stable than O. When O loses the one electron, it enjoys the same orbital stability as N.

Trends in Electron Affinity

21. Define *electron affinity*.

It is a measurement of the degree to which atoms have an attraction for electrons.

22. Which element has the greatest affinity for electrons? _Fluorine__

Trends in Electronegativity

23. Disregarding the noble gases, electronegativity _increases_ as you move across a period, and _decreases_ as you move down a group.

Comprehension Check (from video)

1. List the following atoms in order of increasing atomic radius:

Si, Kr, Cl, K, Ca   Cl, Si, Kr, Ca, K

2. Choose the larger atom in each pair:

Br and **Br$^-$**   **Cl$^-$** and Ar   **K** and K$^+$

Chapter Review (from text)

13. Write the electron configuration of these elements, using a periodic table only.

a. the inert gas in period 3 $1s^2 2s^2 2p^6 3s^2 3p^6$

b. the element in Group 4A, period 4 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

c. the element in Group 2A, period 6 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

20. Indicate which element in each pair has the greater first ionization energy, by circling it.

a. lithium, **boron**   b. **magnesium**, strontium   c. cesium, **aluminum**

Briefly define the following, indicating which outermost sublevels are filled or partially filled (p. 394, 395 in text).

Noble gases – elements in which the outermost s and p sublevels are filled completely.

Representative elements – the outermost s and p sublevels are only partially filled.

Transition metals – the outermost s sublevel and nearby d sublevel contains electrons.

Inner transition metals – the outermost s sublevel and nearby f sublevel contains electrons.